Redox Reactions

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What is an oxidation state?

The oxidation state is a hypothetical charge of an element IF it were 100% ionic.

Take-home message: Treat oxidation number/state like a charge.

Some general rules:

1. Pure elements have oxidation number = 0 — Ex) Br_2: each Br is 0; Ex) Mg: Mg is 0
2. Hydrogen usually has an oxidation number = +1 — Ex) H_2O; CH_4
3. Oxygen usually has an oxidation number = –1 — Ex) H_2O; CO_2
4. Everything else: follow the rules for ionic charges! — Ex) NaCl: Na = +1; Cl = –1; Ex) CF_4: C = +4; F = –1; Ex) CO_2: C = +4; O = –2

Some reactions involve a change in the oxidation state of an atom/element!

What is a redox reaction?

A reaction in which the oxidation states of some elements change.

There are two types:

1. Reduction: Gain of an electron; becomes more negatively charged
2. Oxidation: Loss of an electron; becomes less negatively charged

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$$Ag^{+}\ (aq) + 1e^{-} \rightarrow Ag\ (s)$$

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RIG = Reduction Is Gaining

Choose whatever mnemonic works best for you!

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These are called **half-reactions** because we typically have a both reactions occurring, or a **red**uction-**ox**idation (**redox**) reaction.

How do I balance redox reactions?

$$Fe^{2+}\ (aq) + MnO_4^{-}\ (aq) \rightarrow Fe^{3+}\ (aq) + Mn^{2+}\ (aq)$$

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Fe^{2+} = +2 Mn in MnO_4^-: +7 O in MnO_4^-: –2

Fe^{3+} = +3 Mn^{2+}: +2

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$$\mathbf{Fe^{2+}\,(\mathit{aq}) + MnO_4^{-}\,(\mathit{aq}) \rightarrow Fe^{3+}\,(\mathit{aq}) + Mn^{2+}\,(\mathit{aq})}$$

0. Assign the oxidation states of each element.
1. Separate into half-reactions.

Fe^{2+} = +2 | Mn in MnO_4^-: +7 | O in MnO_4^-: –2

Fe^{3+} = +3 | Mn^{2+}: +2

Ox	1	Fe^{2+} (*aq*)	→	1	Fe^{3+} (*aq*)
Red	1	MnO_4^- (*aq*)	→	1	Mn^{2+} (*aq*)
Overall					

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2. Balance atoms (except O and H): *already balanced*

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2. Balance atoms (except O and H): *already balanced*
3. Balance O atoms with H_2O on opposite side.

Ox	1 $Fe^{2+}\,(aq)$	$\rightarrow$	1 $Fe^{3+}\,(aq)$		
Red	1 $MnO_4^-\,(aq)$	$\rightarrow$	1 $Mn^{2+}\,(aq)$	+	4 $H_2O\,(l)$
Overall					

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Fe^{2+} (*aq*) + MnO_4^- (*aq*) → Fe^{3+} (*aq*) + Mn^{2+} (*aq*)

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4. Balance H atoms with H^+ on opposite side.

Ox				1	Fe^{2+} (*aq*)	→	1	Fe^{3+} (*aq*)			
Red	8	H^+ (*aq*)	+	1	MnO_4^- (*aq*)	→	1	Mn^{2+} (*aq*)	+	4	H_2O (*l*)
Overall											

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1. Separate into half-reactions. $Fe^{3+} = +3$ Mn^{2+}: +2
2. Balance atoms (except O and H): *already balanced*
3. Balance O atoms with H_2O on opposite side.
4. Balance H atoms with H^+ on opposite side.
5. Balance *total charge* with electrons (*e*–).

Ox				1	Fe^{2+} (*aq*)	$\rightarrow$	1	Fe^{3+} (*aq*)			
Red	8	H^+ (*aq*)	+	1	MnO_4^- (*aq*)	$\rightarrow$	1	Mn^{2+} (*aq*)	+	4	H_2O (*l*)
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5. Balance *total charge* with electrons (*e*–).

For the oxidation reaction: the reactants have a total charge of 2+ and the products have a total charge of 3+, so we need 1 extra electron on the product side.

Ox				1	Fe^{2+} (*aq*)	→	1	Fe^{3+} (*aq*)	+	1	e^-
Red	8	H^+ (*aq*)	+	1	MnO_4^- (*aq*)	→	1	Mn^{2+} (*aq*)	+	4	H_2O (*l*)
Overall											

How do I balance redox reactions?

Fe^{2+} (*aq*) + MnO_4^- (*aq*) → Fe^{3+} (*aq*) + Mn^{2+} (*aq*)

0. Assign the oxidation states of each element. $Fe^{2+} = +2$ Mn in MnO_4^-: +7 O in MnO_4^-: –2
1. Separate into half-reactions. $Fe^{3+} = +3$ Mn^{2+}: +2
2. Balance atoms (except O and H): *already balanced*
3. Balance O atoms with H_2O on opposite side.
4. Balance H atoms with H^+ on opposite side.
5. Balance *total charge* with electrons (*e*–).

For the reduction reaction: the reactants have a total charge of 7+ and the products have a total charge of 2+, so we need 5 extra electron on the reactants side.

Ox						1	Fe^{2+} (*aq*)	→	1	Fe^{3+} (*aq*)	+	1	e^-
Red	5	e^-	+	8	H^+ (*aq*)	+ 1	MnO_4^- (*aq*)	→	1	Mn^{2+} (*aq*)	+	4	H_2O (*l*)
Overall													

How do I balance redox reactions?

Fe^{2+} (*aq*) + MnO_4^- (*aq*) → Fe^{3+} (*aq*) + Mn^{2+} (*aq*)

0. Assign the oxidation states of each element.	Fe^{2+} = +2	Mn in MnO_4^-: +7	O in MnO_4^-: –2
1. Separate into half-reactions.	Fe^{3+} = +3	Mn^{2+}: +2	

2. Balance atoms (except O and H): *already balanced*
3. Balance O atoms with H_2O on opposite side.
4. Balance H atoms with H^+ on opposite side.
5. Balance *total charge* with electrons (*e*–).

Ox							1	Fe^{2+} (*aq*)	→	1	Fe^{3+} (*aq*)	+	1	e^-
Red	5	e^-	+	8	H^+ (*aq*)	+	1	MnO_4^- (*aq*)	→	1	Mn^{2+} (*aq*)	+	4	H_2O (*l*)
Overall														

How do I balance redox reactions?

$$\mathbf{Fe^{2+}\ (\mathit{aq}) + MnO_4^-\ (\mathit{aq}) \rightarrow Fe^{3+}\ (\mathit{aq}) + Mn^{2+}\ (\mathit{aq})}$$

0. Assign the oxidation states of each element. Fe^{2+} = +2 Mn in MnO_4^-: +7 O in MnO_4^-: –2
1. Separate into half-reactions. Fe^{3+} = +3 Mn^{2+}: +2
2. Balance atoms (except O and H): *already balanced*
3. Balance O atoms with H_2O on opposite side.
4. Balance H atoms with H^+ on opposite side.
5. Balance *total charge* with electrons (*e*–).
6. Balance the electrons by multiplying entire half-reactions.

Ox			5 × [	1 Fe^{2+} (*aq*)	→	1 Fe^{3+} (*aq*)	+ 1 e^-]
Red	5 e^-	+ 8 H^+ (*aq*)	+	1 MnO_4^- (*aq*)	→	1 Mn^{2+} (*aq*)	+ 4 H_2O (*l*)
Overall							

How do I balance redox reactions?

Fe^{2+} (*aq*) + MnO_4^- (*aq*) → Fe^{3+} (*aq*) + Mn^{2+} (*aq*)

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1. Separate into half-reactions. Fe^{3+} = +3 Mn^{2+}: +2
2. Balance atoms (except O and H): *already balanced*
3. Balance O atoms with H_2O on opposite side.
4. Balance H atoms with H^+ on opposite side.
5. Balance *total charge* with electrons (*e*–).
6. Balance the electrons by multiplying entire half-reactions.

Ox							5	Fe^{2+} (*aq*)	→	5	Fe^{3+} (*aq*)	+	5	e^-		
Red	5	e^-	+	8	H^+ (*aq*)	+	1	MnO_4^- (*aq*)	→	1	Mn^{2+} (*aq*)	+	4	H_2O (*l*)		
Overall																

How do I balance redox reactions?

$$\mathbf{Fe^{2+}\ (\mathit{aq}) + MnO_4^{-}\ (\mathit{aq}) \rightarrow Fe^{3+}\ (\mathit{aq}) + Mn^{2+}\ (\mathit{aq})}$$

0. Assign the oxidation states of each element. $Fe^{2+} = +2$ Mn in MnO_4^-: +7 O in MnO_4^-: –2
1. Separate into half-reactions. $Fe^{3+} = +3$ Mn^{2+}: +2
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3. Balance O atoms with H_2O on opposite side.
4. Balance H atoms with H^+ on opposite side.
5. Balance *total charge* with electrons (*e*–).
6. Balance the electrons by multiplying entire half-reactions.
7. Add the two half-reactions together. (*Make sure overall equation is balanced and no extra electrons).*

Ox							5	$Fe^{2+}\ (aq)$	$\rightarrow$	5	$Fe^{3+}\ (aq)$	+	5	$\cancel{e^-}$
Red	5	$\cancel{e^-}$	+	8	$H^+\ (aq)$	+	1	$MnO_4^-\ (aq)$	$\rightarrow$	1	$Mn^{2+}\ (aq)$	+	4	$H_2O\ (l)$
Overall														

How do I balance redox reactions?

$$\mathbf{Fe^{2+}\ (\mathit{aq}) + MnO_4^-\ (\mathit{aq}) \rightarrow Fe^{3+}\ (\mathit{aq}) + Mn^{2+}\ (\mathit{aq})}$$

0. Assign the oxidation states of each element. $Fe^{2+} = +2$ Mn in MnO_4^-: +7 O in MnO_4^-: –2
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Ox	$5\ Fe^{2+}\ (aq)$	$\rightarrow$	$5\ Fe^{3+}\ (aq) + \cancel{5\ e^-}$
Red	$\cancel{5\ e^-} + 8\ H^+\ (aq) + 1\ MnO_4^-\ (aq)$	$\rightarrow$	$1\ Mn^{2+}\ (aq) + 4\ H_2O\ (l)$
Overall	$8\ H^+\ (aq) + 1\ MnO_4^-\ (aq) + 5\ Fe^{2+}\ (aq)$	$\rightarrow$	$5\ Fe^{3+}\ (aq) + 1\ Mn^{2+}\ (aq) + 4\ H_2O\ (l)$