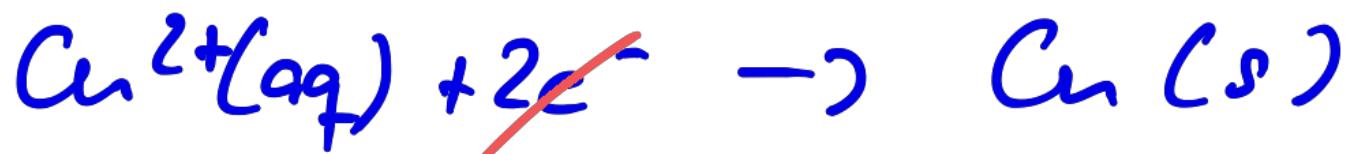


Oxidation Number

Elemental Form	zero (0). Only one kind of atom present, no charge.
Atomic Ions	= the charge on the atom (monoatomic ion)
Group 1A Li, Na, K, Rb, Cs	+1 unless elemental form
Group 2A Be, Mg, Ca, Sr, Ba	+2 unless elemental form
Hydrogen (H)	+1 when bonded to a non-metal, -1 when bonded to a metal (rare)
Oxygen (O)	-1 in peroxides O ₂ ⁻ -2 in all other compounds (most common)
Fluorine (F)	-1, always
Neutral Compounds	The sum of all oxidation numbers of atoms or ions in a neutral compound is zero
Ionic Compounds	The sum of all oxidation numbers of atoms in an ionic compound is the charge on the polyatomic ion.

Oxidation-Reduction Reactions





oxidation: loss of e^- ; ox.-number \uparrow

reduction: gain of e^- ; ox.-number \downarrow

reducing agent: species that is oxidized

oxidizing agent: species that is reduced

Activity Series

Metal	Oxidation Reaction
Lithium	$\text{Li(s)} \rightarrow \text{Li}^{+}(aq) + \text{e}^{-}$
Potassium	$\text{K(s)} \rightarrow \text{K}^{+}(aq) + \text{e}^{-}$
Barium	$\text{Ba(s)} \rightarrow \text{Ba}^{2+}(aq) + 2\text{e}^{-}$
Calcium	$\text{Ca(s)} \rightarrow \text{Ca}^{2+}(aq) + 2\text{e}^{-}$
Sodium	$\text{Na(s)} \rightarrow \text{Na}^{+}(aq) + \text{e}^{-}$
Magnesium	$\text{Mg(s)} \rightarrow \text{Mg}^{2+}(aq) + 2\text{e}^{-}$
Aluminum	$\text{Al(s)} \rightarrow \text{Al}^{3+}(aq) + 3\text{e}^{-}$
Manganese	$\text{Mn(s)} \rightarrow \text{Mn}^{2+}(aq) + 2\text{e}^{-}$
Zinc	$\text{Zn(s)} \rightarrow \text{Zn}^{2+}(aq) + 2\text{e}^{-}$
Chromium	$\text{Cr(s)} \rightarrow \text{Cr}^{3+}(aq) + 3\text{e}^{-}$
Iron	$\text{Fe(s)} \rightarrow \text{Fe}^{2+}(aq) + 2\text{e}^{-}$
Cobalt	$\text{Co(s)} \rightarrow \text{Co}^{2+}(aq) + 2\text{e}^{-}$
Nickel	$\text{Ni(s)} \rightarrow \text{Ni}^{2+}(aq) + 2\text{e}^{-}$
Tin	$\text{Sn(s)} \rightarrow \text{Sn}^{2+}(aq) + 2\text{e}^{-}$
Lead	$\text{Pb(s)} \rightarrow \text{Pb}^{2+}(aq) + 2\text{e}^{-}$
Hydrogen	$\text{H}_2(g) \rightarrow 2\text{H}^{+}(aq) + 2\text{e}^{-}$
Copper	$\text{Cu(s)} \rightarrow \text{Cu}^{2+}(aq) + 2\text{e}^{-}$
Silver	$\text{Ag(s)} \rightarrow \text{Ag}^{+}(aq) + \text{e}^{-}$
Mercury	$\text{Hg(l)} \rightarrow \text{Hg}^{2+}(aq) + 2\text{e}^{-}$
Platinum	$\text{Pt(s)} \rightarrow \text{Pt}^{2+}(aq) + 2\text{e}^{-}$
Gold	$\text{Au(s)} \rightarrow \text{Au}^{3+}(aq) + 3\text{e}^{-}$



 Ease of oxidation increases



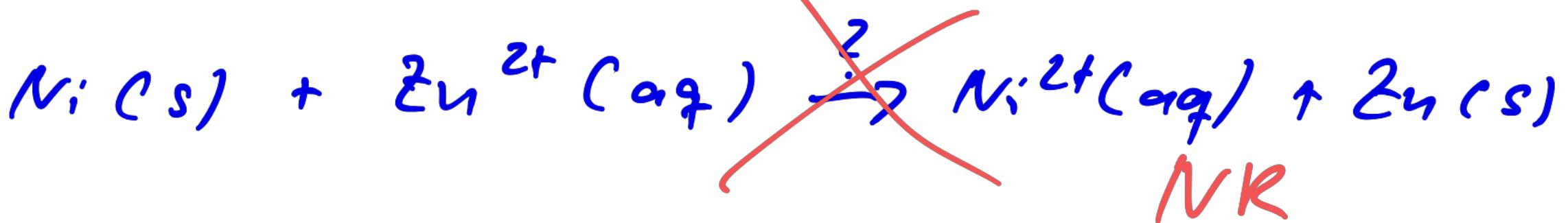
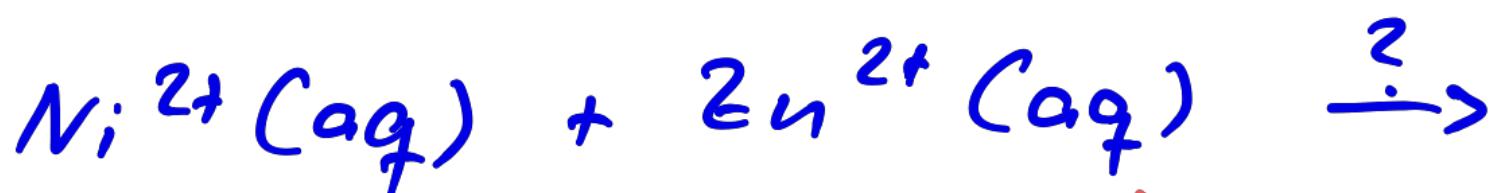
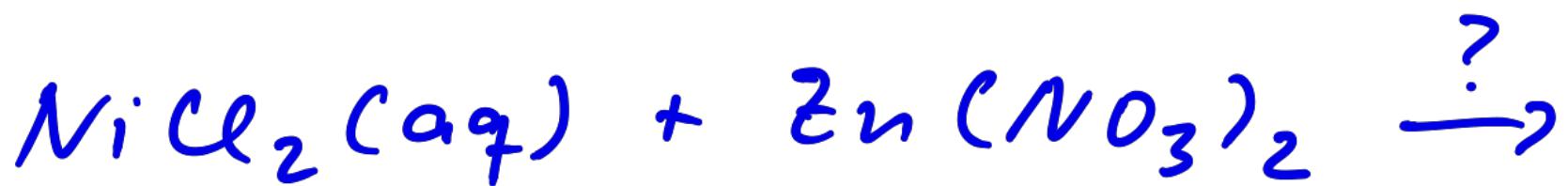
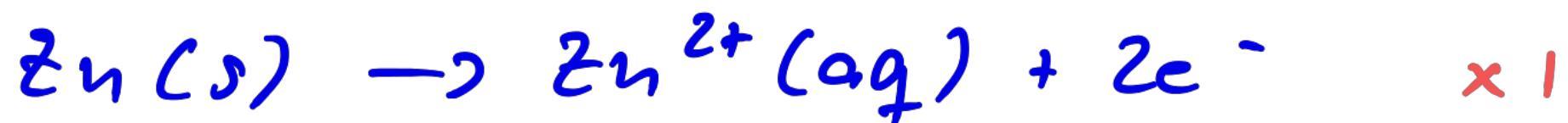
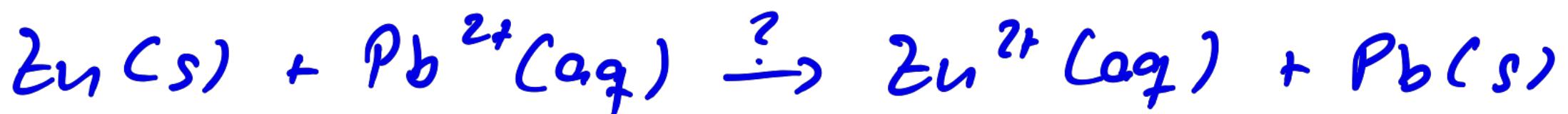
Activity Series

Metal	Oxidation Reaction
Lithium	$\text{Li(s)} \rightarrow \text{Li}^+(aq) + \text{e}^-$
Potassium	$\text{K(s)} \rightarrow \text{K}^+(aq) + \text{e}^-$
Barium	$\text{Ba(s)} \rightarrow \text{Ba}^{2+}(aq) + 2\text{e}^-$
Calcium	$\text{Ca(s)} \rightarrow \text{Ca}^{2+}(aq) + 2\text{e}^-$
Sodium	$\text{Na(s)} \rightarrow \text{Na}^+(aq) + \text{e}^-$
Magnesium	$\text{Mg(s)} \rightarrow \text{Mg}^{2+}(aq) + 2\text{e}^-$
Aluminum	$\text{Al(s)} \rightarrow \text{Al}^{3+}(aq) + 3\text{e}^-$
Manganese	$\text{Mn(s)} \rightarrow \text{Mn}^{2+}(aq) + 2\text{e}^-$
Zinc	$\text{Zn(s)} \rightarrow \text{Zn}^{2+}(aq) + 2\text{e}^-$
Chromium	$\text{Cr(s)} \rightarrow \text{Cr}^{3+}(aq) + 3\text{e}^-$
Iron	$\text{Fe(s)} \rightarrow \text{Fe}^{2+}(aq) + 2\text{e}^-$
Cobalt	$\text{Co(s)} \rightarrow \text{Co}^{2+}(aq) + 2\text{e}^-$
Nickel	$\text{Ni(s)} \rightarrow \text{Ni}^{2+}(aq) + 2\text{e}^-$
Tin	$\text{Sn(s)} \rightarrow \text{Sn}^{2+}(aq) + 2\text{e}^-$
Lead	$\text{Pb(s)} \rightarrow \text{Pb}^{2+}(aq) + 2\text{e}^-$
Hydrogen	$\text{H}_2\text{(g)} \rightarrow 2\text{H}^+(aq) + 2\text{e}^-$
Copper	$\text{Cu(s)} \rightarrow \text{Cu}^{2+}(aq) + 2\text{e}^-$
Silver	$\text{Ag(s)} \rightarrow \text{Ag}^+(aq) + \text{e}^-$
Mercury	$\text{Hg(l)} \rightarrow \text{Hg}^{2+}(aq) + 2\text{e}^-$
Platinum	$\text{Pt(s)} \rightarrow \text{Pt}^{2+}(aq) + 2\text{e}^-$
Gold	$\text{Au(s)} \rightarrow \text{Au}^{3+}(aq) + 3\text{e}^-$

Ease of oxidation increases

tin + hydrochloric acid





Balancing Redox Equations in acidic solutions

Unbalanced:

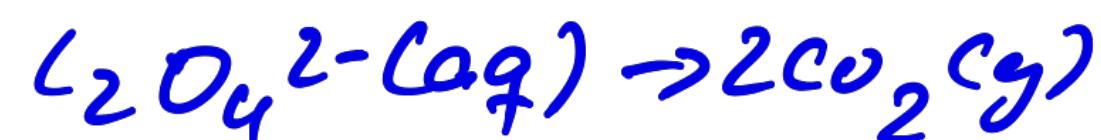


1. Divide into half reactions



2. Balance half reactions

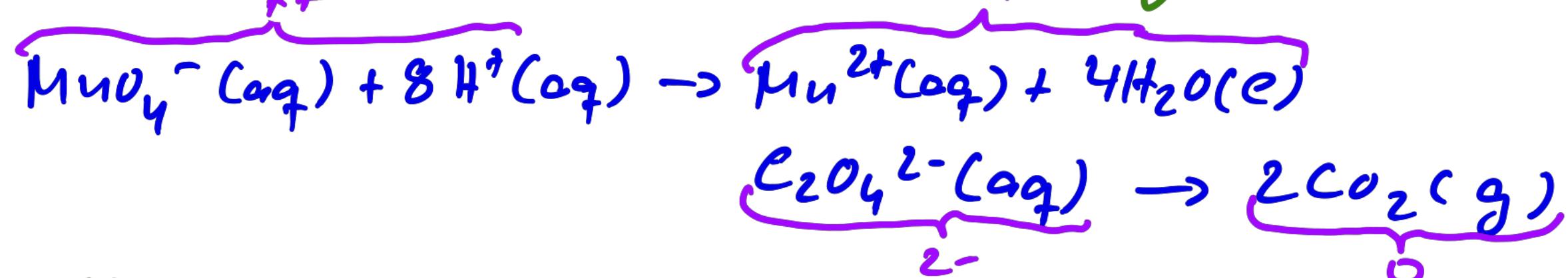
a). Balance elements other than "O" and "H"



b). Balance "O" atoms by adding H_2O as needed



c). Balance "H" atoms by adding H^+ as needed



d). Balance charges by adding " e^- "



3. Equalize e^- for the two half reactions



Balancing Redox Reactions in Basic Solution



initially the same as for acidic solution

1. Half-Reactions

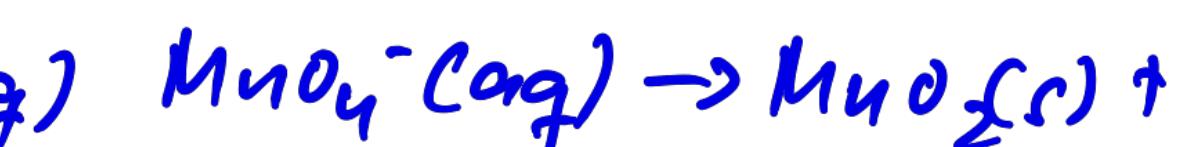


2. Balance Half Reactions

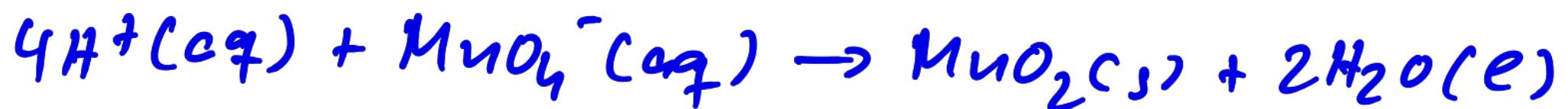
a). Balance elements other than "O" and "H"



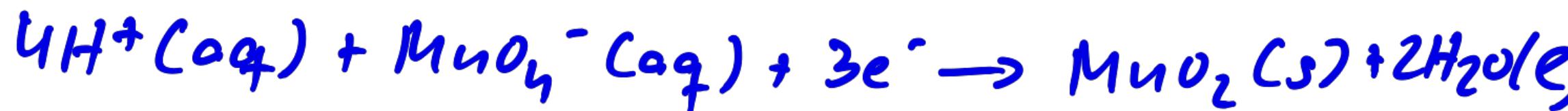
b). balance "O" by adding H_2O



c). Balance "H" by adding H^+



d). Balance charges :



3. Neutralize H^+ by adding OH^- to both sides

