for example: let's say that Fe^{2+} ions are oxidized into Fe^{3+} by dichromate ions ($Cr_2O_7^{2-}$), which are converted into Cr^{3+}.

$$\boxed{Fe^{2+} + Cr_2O_7^{2-} \longrightarrow Fe^{3+} + Cr^{3+}}$$

skeleton rxn.

- often carried out under v. high / low pH
  OH$^-$ H$^+$

let's balance above rxn under acidic conditions.

4/24/15 4-step process to balancing any redox eq$^{\underline{n}}$.

Step(1): Write out unbalanced redox eq. in ionic form.

$$Fe^{2+} + Cr_2O_7^{2-} \longrightarrow Fe^{3+} + Cr^{3+}$$

Step(2): Separate into two half-rxns.

$$\overset{(+2)}{Fe^{2+}} \longrightarrow \overset{(+3)}{Fe^{3+}} \quad (ox)$$

$$\overset{(+6)}{Cr_2O_7^{2-}} \longrightarrow \overset{(+3)}{Cr^{3+}} \quad (red)$$

Step(3): Balance $\frac{1}{2}$-rxns using:

i) Coefficients

* ii) H_2O molecules to balance O

* iii) H^+ ions to balance H

iv) e^- to balance charge.

step(4) Add $\frac{1}{2}$-rxns in such a way as to cancel e's. → Overall eq.

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3) ox $\boxed{Fe^{2+} \longrightarrow Fe^{3+} + e^-}$ i) ✓ ii) ✓ iii) ✓ iv) ?

++ +++ −

RED $6e^- + 14H^+ + Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$ i) ii) iii) iv)

$\underbrace{\qquad\qquad}_{12+}$ $\underbrace{\qquad\qquad}_{6+}$

4) $(Fe^{2+} \longrightarrow Fe^{3+} + e^-) \times 6$

$\oplus$ $(6e^- + 14H^+ + Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O) \times 1$

$$6Fe^{2+} + 14H^+ + Cr_2O_7^{2-} \xrightarrow{6e^-} 6Fe^{3+} + 2Cr^{3+} + 7H_2O$$

(common to leave off state symbols)

For reactions carried out under BASIC (high pH) conditions...

Balance by same method, except we add OH^- ions @ end to neut. H^+.

$$OH^- + H^+ \longrightarrow H_2O$$

last rxn ~ change to BASIC conditions...

$$6Fe^{2+} + \cancel{14H^+} + Cr_2O_7^{2-} \xrightarrow{6e^-} 6Fe^{3+} + 2Cr^{3+} + \cancel{7H_2O}$$

$+ 14OH^-$ $\qquad$ $+ 14OH^-$

$\cancel{14}\ H_2O$ $\quad -7$

$+7H_2O$

so...

$$6Fe^{2+} + 7H_2O + Cr_2O_7^{2-} \xrightarrow{6e^-} 6Fe^{3+} + 2Cr^{3+} + 14OH^-$$

__Galvanic Cells__ – a way to generate elec via a chem rxn.

If we add $Zn(s)$ to a sol^n containing $Cu^{2+}_{(aq)}$ ions, we get a redox rxn:

$$Zn(s) + Cu^{2+}(aq) \xrightarrow{2e^-} Zn^{2+}_{(aq)} + Cu(s)$$

e^-s transfer directly between $Zn + Cu^{2+}$

– hard to generate elec!

– because e^-s transfer: $Zn \xrightarrow{2e^-} Cu^{2+}$

– need to separate out Zn and Cu^{2+}

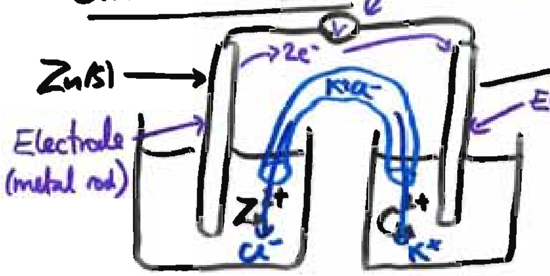
– ½-rxns!

$Zn \longrightarrow Zn^{2+} + 2e^-$ (ox)

$2e^- + Cu^{2+} \longrightarrow Cu$ (red).

Galvanic Cell

voltmeter ← measures voltage / potential difference / electromotive force (emf)



– VOLTAGE

~ Electrical "pressure" (analogy)

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

Oxidation occurs @ ANODE

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

reduction occurs @ CATHODE

RC – COLA

PROBLEM

LHS
- making +ve ions
- makes solⁿ +ve.
- repels more +ve ions from forming.

RHS
- remove +ve ions
- makes solⁿ –ve.
- pulls back + stops Cu^{2+} ions from leaving solⁿ

SOLN

Salt bridge

Cell Diagram Notation:

ANODE (LHS) CATHODE (RHS)

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

electrodes are @ ends

| = phase boundary.

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

Zn-elec (s) Zn^{2+}(aq)

salt bridge

each end = phase boundary.

liq solid