TUESDAY, APRIL 14, 2026 AP CHEMISTRY

CH. 20 ELECTROCHEMISTRY

OXIDATION NUMBERS

SEE ALSO CH. 4 pg 140

OX. NOS ARE A BOOKKEEPING METHOD FOR KEEPING TRACK OF HOW ELECTRONS ARE EXCHANGED IN CHEMICAL REACTIONS.

RULES: ① ELEMENTS AS PURE SUBSTANCES IN THEIR STANDARD FORMS HAVE AN OX. NO. OF ZERO:

Fe Ag Ca Pb Sn Cu

H_2 N_2 O_2 F_2 Cl_2 Br_2 I_2 P_4 S_8

② THE OX. NO. OF A MONATOMIC ION IS EQUAL TO THE ION'S CHARGE:

Na^+ IS +1 Fe^{2+} IS +2

Cl^- IS -1 P^{-3} IS -3

③ NON-METALS FREQUENTLY HAVE NEGATIVE OX. NO.S IN COMPOUNDS

SOME SPECIFIC, USEFUL FACTS ABOUT COMMON ELEMENTS

F IS ALWAYS -1 IN ALL COMPOUNDS

H IS +1 IN COMBINATION WITH NON-METALS
IS -1 IN COMB. WITH METALS

O IS -2 IN NEARLY ALL COMPOUNDS

Cl, Br, I IN BINARY COMPOUNDS ARE -1
AS CENTRAL ATOMS IN OXYANIONS +1 +3 +5 OR +7

④ THE SUM OF OX. NOS. IN FORMULA EQUALS THE OVERALL CHARGE.

TRY THESE:

H_2S

ClO_4^-

Cl_2

H_2SO_4

SO_3^{2-}

$KMnO_4$

MnO_2

ANSWERS:

H_2S H +1, S −2

ClO_4^- Cl +7, O −2

Cl_2 Cl 0

H_2SO_4 H +1, O −2, S +6

SO_3^{2-} S +4, O −2

$KMnO_4$ K +1, Mn +7, O −2

MnO_2 Mn +4, O −2

AN OXIDATION-REDUCTION, OR REDOX, RXN IS ONE IN WHICH ELECTRONS ARE TRANSFERRED BETWEEN SPECIES.

FOR EX. $$2\underset{0}{Na} + \underset{0}{Cl_2} \rightarrow 2\underset{+1}{Na}\underset{-1}{Cl}$$ SYNTHESIS

$$2\underset{+1}{Ag}\underset{+5}{N}\underset{-2}{O_3} + \underset{0}{Cu} \rightarrow 2\underset{0}{Ag} + \underset{+2}{Cu}(\underset{+5}{N}\underset{-2}{O_3})_2$$ SINGLE REPLACEMENT

<u>REDUCTION</u> IS THE GAIN OF ELECTRONS SO THAT OX. NO. IS REDUCED, OR DECREASES:

$$Fe^{3+} + 3e^- \rightarrow Fe$$ METAL REDUCTION $(+) \rightarrow (0)$

$$Cl_2 + 2e^- \rightarrow 2Cl^-$$ NON-METAL REDUCTION $(0) \rightarrow (-)$

TO JUSTIFY THE NAME, CONSIDER THE REDUCTION IN MASS FOR AN ORE TRANSFORMED INTO A METAL:

$$2\underset{+3}{Fe_2}\underset{-2}{O_3} + 3\underset{0}{C} \rightarrow 4\underset{0}{Fe} + 3\underset{+4}{C}\underset{-2}{O_2}$$

$2(159.7) = \boxed{319.4}$ $\qquad$ $4(55.845) = \boxed{223.4}$

Tu 2026-04-14 AP Chem

Oxidation is Loss of electrons so that ox. no. increases:

$$Fe \rightarrow Fe^{3+} + 3e^-$$

Metal oxidation
$(0) \rightarrow (+)$

$$2O^{2-} \rightarrow O_2 + 2e^-$$

Non-metal oxidation
$(-) \rightarrow (0)$ or $(0) \rightarrow (+)$
sometimes

Remember OIL RIG
Oxidation Is Loss of e^-'s
Reduction Is Gain of e^-'s

Be able to analyze a chemical equation to identify what is oxidized and what is reduced:

$$\underset{0}{Zn} + 2\underset{+1}{H}\underset{-1}{Cl} \rightarrow \underset{+2}{Zn}\underset{-1}{Cl_2} + \underset{0}{H_2}$$

Zn is oxidized $Zn \rightarrow Zn^{2+} + 2e^-$

H^+ is reduced $2H^+ + 2e^- \rightarrow H_2$

Terminology in your textbook, but not on the AP Exam:

Zn is a reducing agent b/c it causes H^+ to be reduced.

H^+ is a oxidizing agent b/c it causes Zn to be oxidized.

Demo: Steel wool + hydrochloric acid — note bubbles and color change. Write a chemical equation. What was reduced, what oxidized?

$$\boxed{Fe + 2HCl \rightarrow FeCl_2 + H_2}$$

BREAKING REDOX RXNS INTO TWO SEPARATE BUT LINKED PROCESSES CAN BE TAKEN ONE STEP FURTHER BY WRITING THEM AS HALF-REACTIONS. A TABLE OF THESE IS ON pg871. THE TABLE IS IN ORDER FROM MOST-EASILY-REDUCED TO LEAST AND ALL HALF-REACTIONS ARE WRITTEN AS A REDUCTION. WRITTEN IN REVERSE, THEY BECOME OXIDATION HALF-REACTIONS.

EACH REDUCTION HALF-REACTION INCLUDES THE HALF-CELL REDUCTION POTENTIAL (E°_{red}). WHEN HALF-REACTIONS ARE COMBINED E°_{red} VALUES CAN BE USED TO CALC. THE VOLTAGE OF A BATTERY BASED ON THAT COMBINATION OF HALF-RXNS.

THERMODYNAMICALLY FAVORED COMBINATIONS HAVE A POSITIVE VALUE FOR THE SUM OF E°_{red} AND E°_{ox}, WHERE $E^\circ_{ox} = -E^\circ_{red}$ FOR THE REVERSED RED. RXN.

$$E^\circ_{cell} = E^\circ_{red} + E^\circ_{ox}$$

FOR EX. $Fe^{3+} + 1e^- \rightarrow Fe^{2+} \quad E^\circ_{red} = +0.77\,V$

$Sn^{4+} + 2e^- \rightarrow Sn^{2+} \quad E^\circ_{red} = +0.15\,V$

WRITE THE LOWER-VALUED HALF-RXN IN REVERSE:

$$Sn^{2+} \rightarrow Sn^{4+} + 2e^- \quad E^\circ_{ox} = -0.15\,V$$

ADD THE OX. AND RED. HALF-RXNS, MAKING SURE TO MULT. BY COEFF. AS NECESSARY TO MAKE THE NO. OF e^-S EQUAL:

$$\begin{array}{l} 2Fe^{3+} + \cancel{2e^-} \rightarrow 2Fe^{2+} \\ +\ Sn^{2+} \rightarrow Sn^{4+} + \cancel{2e^-} \\ \hline 2Fe^{3+} + Sn^{2+} \rightarrow 2Fe^{2+} + Sn^{4+} \end{array}$$

$E_{cell} = (+0.77) + (-0.15)$

$E_{cell} = +0.62\,V$

Tu 2026-04-14 AP Chem

E° IS AN INTENSIVE PROPERTY AND ITS VALUE DOES <u>NOT</u> CHANGE WHEN A HALF-RXN IS MULT BY A COEFF.

THE NUMBER OF e^-'S ($n = 2e^-$'S IN OUR EXAMPLE) MUST MATCH SO THEY CANCEL OUT OF THE OVERALL RXN.

THE VALUE OF n IS IMPORTANT IN CALCULATIONS.

THE SIGN OF E° IS REVERSED (+ TO − OR − TO +) WHEN A RXN IS REVERSED.

BALANCED, COMBINED RXNS MUST BALANCE

1. ATOMS
2. CHARGE
3. ELECTRONS

(WE WILL SKIP BALANCING REDOX RXNS BY THE HALF-RXN METHOD IN ACIDIC AND BASIC SOLUTION B/C IT IS COMPLICATED AND NOT TESTED ON THE AP EXAM).

VOLTAIC CELLS

A.K.A., BATTERIES. A VOLTAIC CELL IS A CHEMICAL REACTION COMPARTMENTALIZED SO THAT ELECTRON TRANSFER HAPPENS THROUGH A WIRE. THE TRANSFERRED ELECTRONS CAN BE MADE TO DO WORK.

KEY TERMS

VOLT (V) A MEASURE OF ELECTRICAL POTENTIAL ENERGY AS JOULES (J) PER COULOMB (C) OF CHARGE

$1\,V = 1\,J/C$

VOLTAGE IS INTENSIVE AND B/C IT IS A RATIO IT IS INDEPENDENT OF THE AMOUNT OF MATERIAL.

COULOMB (C): THE UNIT OF ELECTRICAL CHARGE

$1\,e^-$: 1.602×10^{-19} C

$1\,\text{mol}\,e^-$: 96,485 C/mol = F

THE FARADAY CONSTANT

emf = ELECTROMOTIVE FORCE

A SYNONYM FOR VOLTS AND EQUIVALENT TO Ecell, WHICH IS THE VOLTAGE FOR A PAIR OF HALF-RXNS WHEN COMBINED

Ecell IS POSITIVE (+) FOR A SPONTANEOUS RXN.
Ecell IS NEGATIVE (−) FOR A RXN SPON. IN REVERSE.

HOW TO DIAGRAM A VOLTAIC CELL

FIRST, THE CHEMICAL EQUATIONS:

$$Cu^{2+} + 2e^- \rightarrow Cu \quad E^\circ_{red} = +0.34V$$

~~$Zn^{2+} + 2e^- \rightarrow Zn \quad E^\circ_{red} = -0.76V$~~

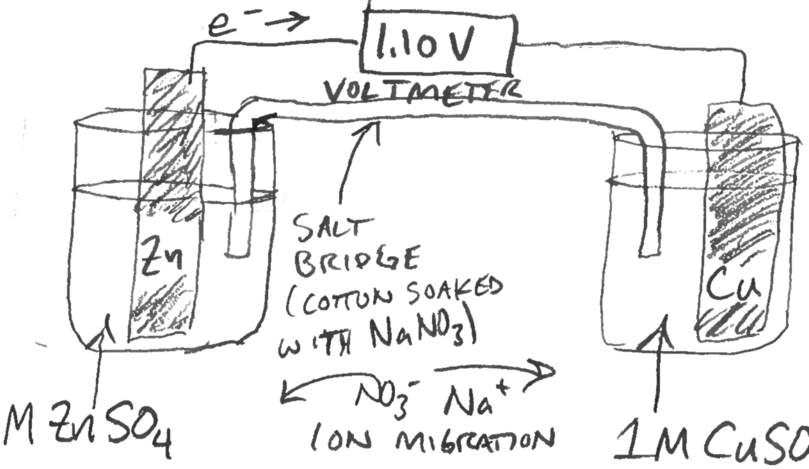
$$+ \quad Zn \rightarrow Zn^{2+} + 2e^- \quad E^\circ_{ox} = +0.76V$$

$$Cu^{2+} + Zn \rightarrow Zn^{2+} + Cu \quad E^\circ_{cell} = +0.34V + (+0.76V) = 1.10V$$

SECOND, THE PHYSICAL SET-UP:

CHEMICALS: Zn
$ZnSO_4$
ANODE
(OXIDATION)

CHEMICALS: Cu
$CuSO_4$
CATHODE
(REDUCTION)



(SALT BRIDGES COMPLETE THE CIRCUIT B/C ELECTROLYTES CONDUCT ELECTRICITY).

NOTE: THE SOLID PIECE OF ZINC LOSES MASS AS THE RXN PROCEEDS B/C THE Zn TURNS INTO SOLUBLE Zn^{2+} IONS. THE SOLID PIECE OF COPPER GAINS MASS B/C THE DISSOLVED Cu^{2+} IONS TURN INTO SOLID Cu.

IN THEORY ANY PAIR OF HALF-REACTIONS (AS TABULATED ON pg 871 AND IN APPENDIX E) CAN BE COMBINED TO MAKE A VOLTAIC CELL. REVERSE THE REDUCTION HALF-RXN WITH THE LOWER HALF-CELL POTENTIAL, EVEN OUT ELECTRONS, AND ADD!

Tu 2026-04-14 APCHEM

NOTE: THE STANDARD HYDROGEN ELECTRODE (S.H.E.) HAS A DEFINED HALF-CELL POTENTIAL OF 0.00V. ALL OTHER HALF-CELL POTENTIALS ARE MEASURED RELATIVE TO THAT.

HOW TO CREATE AN ARBITRARY VOLTAIC CELL:

GIVEN TWO REDUCTION HALF-RXNS:

$$Ni^{2+} + 2e^- \rightarrow Ni \quad E^\circ_{red} = -0.28V$$

$$Al^{3+} + 3e^- \rightarrow Al \quad E^\circ_{red} = -1.66V$$

REVERSE THE ONE WITH A LOWER E°_{red}:

$$Al \rightarrow Al^{3+} + 3e^- \quad E_{ox}^\circ = +1.66V$$

(REVERSE SIGN)

DO THIS SO THAT THE SUM IS POSITIVE:

$$E^\circ_{cell} = E^\circ_{red} + E^\circ_{ox}$$

$$E^\circ_{cell} = \underset{Ni}{(-0.28V)} + \underset{Al}{(+1.66V)} = \boxed{+1.38V}$$

EVEN OUT THE ELECTRONS WITH A MULTIPLIER, AS NEEDED, AND COMBINE TO WRITE THE OVERALL NET IONIC EQUATION:

$$3(Ni^{2+} + 2e^- \rightarrow Ni) = 3Ni^{2+} + \cancel{6e^-} \rightarrow 3Ni$$

$$2(Al \rightarrow Al^{3+} + 3e^-) = + 2Al \rightarrow 2Al^{3+} + \cancel{6e^-}$$

$$3Ni^{2+} + 2Al \rightarrow 2Al^{3+} + 3Ni$$

$n = 6$ mol e^- TRANSFERRED

IN THE CHEM. EQN. THE NO. OF e^-S CANCELS OUT BUT IS IMPORTANT FOR LATER CALCULATIONS.

THE TABLE OF pg 871 REDUCTION POTENTIALS IS A MORE DETAILED VERSION OF THE ACTIVITY SERIES FROM CH. 4. pg 144

TAKEN IN ORDER, FROM MOST POSITIVE THROUGH ZERO TO MOST NEGATIVE, THE TABLE QUANTIFIES THE RELATIVE THERMODYNAMIC FAVORABILITY OF THE REDUCTION OF THE IONS SHOWN. IONS AT THE TOP ARE "EASILY REDUCED" AND SO SERVE AS OXIDIZING AGENTS. IONS AT THE BOTTOM ARE USEFULLY THOUGHT OF AS THE PRODUCTS OF A REVERSED HALF-RXN.

CONSIDER:

$Cu^{2+} + 2e^- \rightarrow Cu \quad E^\circ_{red} = +0.34V$

$2H^+ + 2e^- \rightarrow H_2 \quad E^\circ_{red} = 0.00V$

$Zn^{2+} + 2e^- \rightarrow Zn \quad E^\circ_{red} = -0.76V$

SINCE A COMBINATION OF Cu AND H^+ IONS WOULD GIVE AN E°_{cell} OF $-0.34V$, ACIDS DO NOT DISSOLVE Cu. ON THE OTHER HAND, A COMBINATION OF Zn AND H^+ IONS GIVEN AN E°_{cell} OF $+0.76V$ AND Zn DISSOLVES READILY IN ACID. SEE MY DISSOLVED Cu-PLATED Zn PENNY.

FREE ENERGY AND REDOX RXNS

THERMITE SPHERES $\underset{+3\ -2}{Fe_2O_3} + \underset{0}{2Al} \longrightarrow \underset{+3\ -2}{Al_2O_3} + \underset{0}{2Fe}$

(THIS RXN HAS A VERY LARGE $-\Delta G$)

E_{cell} OR CELL POTENTIAL GIVES INFO. ABOUT THERMODYNAMIC FAVORABILITY (AKA, SPONTANEITY): FOR $+E_{cell}$, A RXN IS SPONTANEOUS

FOR $-E_{cell}$, A RXN IS NON-SPON.

TU 2026-04-14 AP CHEM

E IS RELATED TO G:

$$\Delta G = -nFE \quad \text{OR} \quad \Delta G^\circ = -nFE^\circ$$

FOR STANDARD CONDITIONS

n = THE NO. OF mol OF e^-'S TRANSFERRED

F = FARADAY'S CONSTANT, WHICH IS THE AMOUNT OF CHARGE IN 1 mol OF ELECTRONS: 96,485 C/mol OR $\frac{J}{mol \cdot V}$

FOR THIS EQN, CONVERT kJ TO J FOR ΔG SO THE UNITS WORK OUT:

$$\Delta G = -nFE$$

$$(J) = -(mol)\left(\frac{C}{mol}\right)\left(\frac{J}{C}\right)$$

REMEMBER:
$1V = 1J/C$

SO, GIVEN ΔG, YOU CAN CALCULATE E, AND VICE VERSA. YOU NEED A BALANCED CHEMICAL EQN, WHICH YOU CAN USE TO FIGURE OUT THE VALUE OF n.

IT IS ALSO POSSIBLE TO RELATE E° TO K_{eq}, JUST AS WE RELATED ΔG° TO K_{eq} IN CH. 19:

$$\Delta G^\circ = -RT \ln K_{eq}$$

$$-nFE^\circ = -RT \ln K_{eq} \quad \text{SO} \quad \boxed{E^\circ = \frac{RT}{nF} \ln K_{eq}}$$

$R = 8.314 \frac{J}{K\,mol}$

T IN KELVIN (K)

WHILE WE'RE AT IT, WE CAN ALSO TAKE A LOOK AT HOW CELL VOLTAGE DEPENDS ON Q, THE RXN QUOTIENT:

$$\Delta G = \Delta G^\circ + RT \ln Q$$

$$\text{SO} \quad -nFE = -nFE^\circ + RT \ln Q$$ (DIVIDE BOTH SIDES BY $-nF$)

★ $\boxed{E = E^\circ - \frac{RT}{nF} \ln Q}$ THE NERNST EQUATION

$$E = E^\circ - \frac{RT}{nF}\ln Q \qquad Q = \frac{[\text{PRODUCTS}]}{[\text{REACTANTS}]}$$

FOR THE AP EXAM YOU WILL NOT NECESSARILY NEED TO DO CALCULATIONS WITH THIS EQUATION. HOWEVER, THEY DO REQUIRE YOU TO USE IT TO GUIDE YOUR CONCEPTUAL REASONING FOR QUALITATIVE DESCRIPTIONS.

WE WILL DO SOME CALCULATIONS TO GET A SENSE OF HOW IT WORKS.

$$\boxed{Zn + Cu^{2+} \rightarrow Cu + Zn^{2+} \quad E^\circ = 1.10V}$$

STANDARD CONDITIONS

T=298K ALL CONC. = 1M SO Q = 1

$$E = 1.10V - \frac{RT}{(2)F}\ln 1 = 1.10V$$

UNDER STD. COND., $E = E^\circ$

BATTERY VOLTAGE BOOST AND LIFE EXTENDER

AS THIS REDOX RXN MOVES TOWARD EQUILIBRIUM $[Cu^{2+}]$ WILL DECREASE AND $[Zn^{2+}]$ WILL INCREASE. WHEN Cu^{2+} IS USED UP, THE BATTERY IS "DEAD". SO TO EXTEND THE BATTERY LIFE IT CAN BE STARTED WITH $[Cu^{2+}] > 1M$ AND $[Zn^{2+}] < 1M$. $Q = \frac{[Zn^{2+}]}{[Cu^{2+}]} = \frac{1}{1} = 1$

STD COND

BOOST: $Q = \frac{0.01}{1.2} = 8.33\times10^{-3}$

$$E = E^\circ - \frac{RT}{nF}\ln Q = 1.10V - \frac{(8.314\frac{J}{K\,mol})(298K)}{(2\,mol\,e^-)(96485\frac{C}{mole^-})}\ln 8.33\times10^{-3}$$

$$E = 1.10V - (-0.061V) = +1.16V$$

$\boxed{\text{SO AS } Q \text{ GETS SMALLER, VOLTAGE INCREASES}}$ B/C $\ln Q < 0$ FOR $Q < 1$

THIS IS THE SAME AS SAYING ΔG GOES DOWN IN VALUE WITH GREATER THERMODYNAMIC FAVORABILITY.

Tu 2026-04-14 AP CHEM

BATTERIES RUN DOWN

AS THE REACTANTS IN A BATTERY ARE USED UP, AND AS PRODUCT CONC. INCREASE, A BATTERY'S VOLTAGE DROPS. FUNCTIONALLY, ONCE IT GOES BELOW THE MINIMUM VOLTAGE NECESSARY TO RUN A DEVICE, THE DEVICE STOPS WORKING. WE CALL THAT BATTERY "DEAD" BUT A TRULY "DEAD" BATTERY WOULD HAVE $E = 0\,V$.

$$Zn + Cu^{2+} \rightarrow Cu + Zn^{2+} \qquad E^\circ = 1.10V$$

"DYING" BATTERY: $Q = \dfrac{[Zn^{2+}]}{[Cu^{2+}]} = \dfrac{1.5M}{0.0001M} = 15000$

$$E = 1.10V - \frac{RT}{nF}\ln 15000 = 1.10 - 0.12V = +0.98V$$

SO AS Q <u>INCREASES</u>, VOLTAGE <u>DECREASES</u>.

B/C $\ln Q > 0$ FOR $Q > 1$

OR, AS Q <u>INCR</u>, ΔG GOES UP IN VALUE WITH A <u>LESSER</u> THERMODYNAMIC FAVORIBILITY.

THE EFFECT OF TEMP. CAN'T BE ANALYZED JUST USING $E = E^\circ - \frac{RT}{nF}\ln Q$ B/C E° IS DEPENDENT ON TEMPERATURE: CALC. ΔG° AT A NEW TEMP. USING ΔH° AND ΔS° AND $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$. THEN CALC. A NEW VALUE FOR E° USING $\Delta G^\circ = -nFE^\circ$.

GENERALLY, AS T RISES, E° DECREASES AND AS T FALLS, E° INCREASES IN VALUE.

BECAUSE WE STICK TO 298K FOR THESE CALCULATIONS, AND B/C SOME FOLKS PREFER $\log_{10}$ OVER ln, WE CAN TAKE A SHORTCUT IN THE NERNST EQN.:

$$\boxed{E = E^\circ - \left(\frac{0.0592V}{n}\right)\log Q}$$

BATTERIES AND FUEL CELLS

COMMERCIAL BATTERIES ARE SELF-CONTAINED VOLTAIC CELLS, OR SERIES OF CELLS.

AAA, AA, C, AND D CELL BATTERIES ARE SINGLE CELLS WITH A LABELLED VOLTAGE OF 1.5V. CONNECTED UP PLUS TO MINUS, OR IN SERIES, VOLTAGES ADD UP. BIGGER CELLS JUST HAVE MORE REACTANTS INSIDE SO THEY LAST LONGER.

LANTERN (6V), 9-VOLT, AND 12V BATTERIES HAVE MULTIPLE CELLS IN SERIES INSIDE THEM.

DIFFERENT CHEMICAL RXNS ALLOW FOR DIFFERENT BATTERY FEATURES SUCH AS VOLTAGE, LIFESPAN, AND ENERGY DENSITY. A RECHARGABLE BATTERY USES A REVERSIBLE RXN WHICH DELIVERS POWER WHEN DISCHARGED AND RUNS BACKWARD, REPLENISHING REACTANTS, WHEN ELECTRICITY IN RUN THROUGH IT IN REVERSE.

A FUEL CELL IS LIKE A BATTERY IN THAT IT USES A SEPARATED CHEM. RXN. TO GENERATE ELECTRICITY. THE DIFF. IS THAT IT CAN BE CONTINUOUSLY SUPPLIED WITH FUEL AND IS NOT SELF-CONTAINED. A HYDROGEN FUEL CELL COMBINES H_2 AND O_2 TO MAKE ELECTRICITY AND WATER, FOR EXAMPLE.

CORROSION

MOST METALS WILL REACT WITH O_2 AND H_2O TO FORM OXIDES. MANY FORM IMPERVIOUS AND VERY THIN LAYERS ON THEIR SURFACE, WHICH PREVENTS FURTHER OXIDATION.

Tu 2026-04-14 AP CHEM

IRON CORROSION IS DIFFERENT — AS OXIDATION PROCEEDS MORE AND MORE OF THE METAL IS CONSUMED. THIS IS IN PART DUE TO THE FACT THAT IRON (III) OXIDE IS SO MUCH LESS DENSE THAN IRON: THE MATERIAL EXPANDS AND FLAKES OFF AS IT FORMS.

Fe $D = 7.9 g/cm^3$ Fe_2O_3 $D = 5.2 g/cm^3$

IRON RUSTS IN CONTACT WITH WATER AND AIR. SALTY WATER AIDS THE PROCESS B/C ELECTROLYTES CONDUCT ELECTRICITY. IF $pH > 9$, NO RUSTING CAN TAKE PLACE.

TO PREVENT RUST, TWO PRIMARY METHODS ARE USED:

① PAINT — KEEP O_2 AND H_2O AWAY

② CATHODIC PROTECTION — MAKE AN ELECTRICAL CONNECTION TO A MORE-EASILY OXIDIZED METAL

(CATHODE = REDUCTION
ANODE = OXIDATION)

IF YOU CONNECT Fe TO Zn, OR PLATE THE Zn RIGHT ONTO THE IRON (A PROCESS KNOWN AS GALVINIZATION) YOU MAKE THE Fe INTO A CATHODE AND THE Zn INTO AN ANODE. IN THIS WAY, THE Zn OXIDIZES AND GIVES ITS ELECTRONS TO Fe, KEEPING IT FROM RUSTING.

(SEE ALSO SACRIFICIAL ANODES)

ELECTROLYSIS

REDOX RXNS WHICH ARE SPONTANEOUS IN REVERSE CAN BE DRIVEN IN REVERSE BY AN APPLIED VOLTAGE.

DEMONSTRATIONS: I. ELECTROLYSIS OF WATER (HOFFMANN APPARATUS)

OX.: ANODE $(2H_2O \rightarrow O_2 + 4H^+ + 4e^-)$ $E^\circ = -1.23V$

RED.: CATHODE $+(4H_2O + 4e^- \rightarrow 2H_2 + 4OH^-)$ $E^\circ = -0.83V$

$$6H_2O \rightarrow 2H_2 + O_2 + 4H^+ + 4OH^- \qquad E^\circ_{elec} = -2.06V$$

$$6H_2O \rightarrow 2H_2 + O_2 + 4H_2O$$

$$\boxed{2H_2O \rightarrow 2H_2 + O_2}$$

II. ELECTROPLATING Cu ONTO Ag

NEG. CHARGE TO Cu: REDUCES Cu ONTO SURFACE

POS. CHARGE TO Cu: OXIDIZES Cu OFF OF SURFACE

ELECTROLYSIS IS USED TO PURIFY Cu: $2Cu_2O + C \rightarrow CO_2 + 4Cu$

AND TO REFINE Al FROM Al_2O_3 ECONOMICALLY USING CHEAP HYDROELECTRIC POWER.

WHEN A METAL IS COLLECTED AS A SOLID BY ELECTROLYSIS IT IS CALLED "PLATING OUT".

KEY TERMS: **CURRENT** IS THE QTY. OF CHARGE (C) TRANSFERRED PER UNIT TIME (s). UNITS: $\frac{C}{s}$ OR A

$1\,A = 1\,\frac{C}{s}$ A = AMPERE

POWER IS THE QTY. OF ENERGY (J) PER UNIT TIME (s)

UNITS: $\frac{J}{s} = W$ $1W = 1J/s$ W = WATT

PLATING IS THE PROCESS OF USING ELECTRICAL CURRENT TO GIVE ELECTRONS TO METAL IONS, REDUCING THEM TO THE METALLIC STATE

Tu 2026-04-14 AP CHEM

THE CB LIKES TO PUT QUESTIONS ABOUT ELECTROLYSIS ON THE AP EXAM. THE EXAM REFERENCE INFORMATION INCLUDES THE EQUATION

$$I = \frac{q}{t} \qquad \begin{aligned} I &= \text{CURRENT (A)} \\ q &= \text{CHARGE (C)} \\ t &= \text{TIME (s)} \end{aligned} \qquad \text{SO } q = It \quad \text{OR } t = \frac{q}{I}$$

QUESTIONS YOU MIGHT SEE INCLUDE:

1. ELAPSED TIME TO PLATE OUT A SPECIFIC MASS OF A METAL
2. MASS OF METAL PLATED OUT WITH A GIVEN CURRENT AND TIME INTERVAL
3. (MAYBE) HOW MUCH POWER OR ENERGY WAS USED IN A PLATING PROCESS.

EX. 1 HOW LONG WILL IT TAKE TO PLATE OUT 1.0g OF Cu GIVEN SUFFICIENTLY HIGH VOLTAGE AND A CURRENT OF 1.0A?

FIRST, CONSIDER THE RELEVANT REDUCTION HALF-RXN:

$$Cu^{2+} + 2e^- \rightarrow Cu \qquad \text{SO } n = \frac{2 \text{ mol } e^-}{1 \text{ mol Cu}}$$

REASONING FROM UNITS YOU NEED TO DO THIS:

$$\left(\begin{matrix}\text{MASS OF}\\ \text{METAL}\end{matrix}\right) \times \left(\frac{1 \text{ mol METAL}}{\text{MOLAR MASS}}\right) \times \left(\frac{n \text{ mol } e^-}{1 \text{ mol METAL}}\right) \times \left(\frac{\text{FARADAY'S CONSTANT } C}{\text{mol } e^-}\right) \times \left(\begin{matrix}\text{INVERSE}\\ \text{OF CURRENT}\\ \frac{s}{C}\end{matrix}\right)$$

$$1.0\text{ g} \times \frac{1 \text{ mol Cu}}{63.546\text{g}} \times \underbrace{\frac{2 \text{ mol } e^-}{1 \text{ mol Cu}}}_{\text{(FROM HALF-RXN)}} \times \frac{96{,}485\text{C}}{1 \text{ mol } e^-} \times \underbrace{\frac{1\text{s}}{1.0\text{C}}}_{(1A = 1\frac{C}{s})} \cdot \frac{1 \text{ min}}{60\text{s}} =$$

$$\boxed{51 \text{ min}}$$

FARADAY'S CONSTANT (F) IS THE CHARGE OF A SINGLE ELECTRON (1.602×10^{-19} C) TIMES AVOGADRO'S NUMBER (6.022×10^{23}/mol).

Ex. 2 WHAT MASS OF Al IS PRODUCED USING A 10.0A CURRENT FOR 1 hr FROM AN APPROPRIATE SOLUTION OF Al^{3+} IONS?

HALF-RXN: $Al^{3+} + 3e^- \rightarrow Al$

$$\text{TOTAL TIME} \times \left(\text{CONVERT TO S}\right) \times \left(\text{CURRENT } \frac{C}{S}\right) \times \left(\frac{1}{F}\left(\frac{\text{mol } e^-}{C}\right)\right) \times \left(\frac{\text{mol METAL}}{\text{mol } e^-}\right) \times \left(\text{MOLAR MASS } \frac{g}{\text{mol}}\right)$$

$$1\,hr \times \frac{3600\,s}{1\,hr} \times \frac{10.0\,C}{1\,s} \times \frac{1\,mol\,e^-}{96{,}485\,C} \times \frac{1\,mol\,Al}{3\,mol\,e^-} \times \frac{26.982\,g}{1\,mol\,Al} =$$

$(10.0A = 10\,C/s)$ (FROM HALF-RXN)

$$\boxed{3.4\,g\,Al}$$

Ex. 3 GIVEN THAT THE ELECTROLYSIS OF Al ABOVE WAS CARRIED OUT AT 25V, HOW MUCH ENERGY WAS REQUIRED? ALSO, HOW MUCH POWER? AND, AT HOUSEHOLD RATES, WHAT WOULD THAT COST?

FOR FUN (?)

$$1\,hr \cdot \frac{3600\,s}{1\,hr} \cdot \frac{10.0\,C}{1\,s} \cdot \frac{25\,J}{1\,C} = 900{,}000\,J \text{ OR } 900\,kJ$$

$$\left(1\,V = 1\,\frac{J}{C}\right)$$

POWER IS $\frac{\text{ENERGY}}{\text{TIME}}$ OR $\frac{J}{s}$ SO $\frac{900\,kJ}{3600\,s} = 0.25\,kW$

$$\left(1\,kW = \frac{1\,kJ}{1\,s}\right)$$

POWER IS BILLED PER $kW \cdot h$ (1 kW USED FOR 1 hr IS $1\,kW \cdot h$)

$0.25\,kW \cdot 1\,hr = 0.25\,kW \cdot h$

AT 31¢ PER $kW \cdot h$ THIS IS 7.75¢.