Edvantage Science AP Chemistry 2

Chapter 7

Traffic Light Study Guide

Section	Page	I can ...	Red	Amber	Green
7.1	430	Define *oxidation-reduction/redox reactions* and cite types of reactions that are redox reactions.	○	○	○
	431 - 432	Determine the oxidation number of each atom in a chemical species.	○	○	○
	433 - 434	Define *oxidation* and *reduction* and relate each to an atom's change of oxidation number.	○	○	○
	434 - 436	Define *oxidizing agent* and *reducing agent* and identify each in a redox reaction.	○	○	○
7.2	441 - 442	Define *half-reaction*.	○	○	○
	442 - 444	Balance half-reactions both under acidic and basic conditions.	○	○	○
	445 - 446	Balance redox reactions using the half-reaction method.	○	○	○
	447 - 448	Balance *disproportionation* and *comproportionation* reactions.	○	○	○
7.3	457 - 460	Use the SRP Table to determine whether a spontaneous redox reaction, such as a single replacement reaction, will occur.	○	○	○
	460	Relate the strength of a reducing agent (e.g. A^-) to the strength of its complementary oxidizing agent (e.g. A).	○	○	○
	461	Define *oxidation potential* and *reduction potential* and read these off the SRP Table.	○	○	○
	462 - 463	Determine the predominant redox reaction that will occur in a mixture of oxidizing and reducing agents.	○	○	○
	464	Describe a redox titration and use redox titration data to calculate an analyte's concentration.	○	○	○

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7.4	470 - 472	Draw a *standard electrochemical cell*, label its parts, and describe its operation.	○	○	○
	472	Describe a non-metal electrode.	○	○	○
	472 - 474	Use the SRP Table to predict *standard cell potentials* (E^o).	○	○	○
	475 - 476	State the effect of altering half-cell concentrations on a given cell's potential.	○	○	○
	475 - 476	Describe and explain what happens to a cell's potential as it operates. State the meaning and voltage of a *cell at equilibrium*.	○	○	○
	476 - 479	Use the *Nernst Equation* to determine the potential of a cell with non-standard ion concentrations.	○	○	○
	476 - 479	Use the *Nernst Equation* to convert a cell's E^o into K_{eq} (and vice-versa).	○	○	○
	480 - 482	Describe the basic structure of the alkaline dry cell, the lead-acid storage battery, and the fuel cell. Name the common uses for each.	○	○	○
	482 - 483	Define *corrosion*. Describe the chemistry of rusting. Cite factors that increase corrosion and describe methods (particularly *cathodic protection*) for protecting metals from corrosion.	○	○	○
7.5	492 - 493	Draw an *electrolytic cell*, label its parts, and describe its operation.	○	○	○
	493 - 494	Use the SRP Table to predict the voltage required to operate an electrolytic cell.	○	○	○
	495	Contrast the electrolytic cell with the electrochemical cell.	○	○	○
	496 - 498	Describe the difference between a *type 1*, a *type 2*, and a *type 3*, electrolytic cell.	○	○	○
	496 - 498	Use the SRP Table to determine the oxidation half-reaction and the reduction half-reaction that will occur in any given electrolytic cell.	○	○	○
	500 - 501	Describe the following applications of electrolysis: i. *electrowinning*, *electroplating*, and *electrorefining* ii. the *Héroult-Hall process* for producing aluminum iii. the *chloralkali industry* iv. *impressed current cathodic protection*	○	○	○

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7.6	505 - 506	In an electrochemical or electrolytic cell, given *Faraday's constant (F)* and any two of the following, calculate the third. i. the amount of a chemical produced at an electrode ii. the amperage iii. the time the circuit was running.	○	○	○
	507 - 508	Calculate the Standard Free Energy Change (ΔG^o) accompanying an electrochemical reaction.	○	○	○
	509 - 510	Interconvert the Standard Free Energy Change (ΔG^o), the Standard Cell Potential (E^o), and the Equilibrium Constant (K) for an electrochemical reaction and relate each to the reaction's spontaneity.	○	○	○