

Periodic Table with Chemistry Formulas

sparkcharts

CHEMISTRY FORMULAS

ATOMS AND SUBATOMIC PARTICLES

A. Atomic number: The number of protons in an atom of a given element. In the periodic table, elements are ordered by increasing atomic number.

Ex: ^{12}C $Z =$ atomic number

B. Mass number: The number of protons plus the number of neutrons in an atom of an element, measured in atomic mass units (amu). 1 amu = mass of one proton or neutron = 1.66×10^{-27} kg.

Ex: ^{12}C $12 =$ mass number

Name	Symbol	Mass	Charge
Proton	p^+	1 amu	+1
Neutron	n^0	1 amu	0
Electron	e^-	$1/1837$ amu	-1

C. Isotopes: Various atoms of an element that possess different numbers of neutrons, and therefore different mass numbers.

Ex: Boron has ^{10}B , Boron has ^{11}B

D. Calculating atomic mass: An element's atomic mass number on the periodic table is a weighted average of that element's isotopes.

Ex:

Isotope	Abundance	Mass #	Isotope Mass
Boron-10	20.2%	10	10.0129 amu
Boron-11	79.8%	11	11.0093 amu

Atomic mass = $(20.2)(10.0129 \text{ amu}) + (79.8)(11.0093 \text{ amu})$
= 10.81 amu

MOLECULAR MASS

A. Avogadro's number: 6.022×10^{23}

B. The atomic mass number of an element equals the mass (in grams) of one mole of that element.

Ex: One mole of boron has a mass of 10.81 grams.

MOLECULAR FORMULA AND EMPIRICAL FORMULA

A. Molecular formula: The chemical formula of a compound that gives the actual number of atoms of each element in a molecule of the compound.

Ex: The molecular formula for glucose is $\text{C}_6\text{H}_{12}\text{O}_6$.

B. Empirical formula: A "simplified" expression of a chemical formula.

Ex: The empirical formula for glucose is $\text{C}_2\text{H}_4\text{O}$.

BALANCING CHEMICAL EQUATIONS

1. Check for diatomic molecules such as H_2 and O_2 .

- Balance metals.
- Balance nonmetals.
- Balance oxygen.
- Balance hydrogen.

2. You may change only coefficients, never the formula subscripts! Remember all atoms in the equation have equal numbers on both sides of the equation.

3. Reverse the equation so that coefficients are the smallest whole numbers possible.

PERCENTAGE AND PPM

A. Percentage composition: Tells how much mass any object or element contributes to a compound.

Percentage composition = $\frac{\text{mass of element}}{\text{mass of compound}} \times 100$

B. Parts per million (ppm): In a homogeneous solution, the number of grams of a

particular compound that would be present in a one million gram sample of the compound or solution.

TEMPERATURE CONVERSION

A. Converting between Fahrenheit and Celsius

$$\text{Fahrenheit} = \frac{9}{5}(\text{Celsius}) + 32^\circ\text{F}$$

$$\text{Celsius} = \frac{5}{9}(\text{Fahrenheit} - 32^\circ\text{F})$$

B. Kelvin: The temperature in K, absolute zero, is the theoretical lowest possible temperature, at which all molecular activity ceases.

$$0^\circ\text{K} = -273.15^\circ\text{C} = -459.67^\circ\text{F}$$

$$\text{K} = ^\circ\text{C} + 273.15 \quad ^\circ\text{C} = \text{K} - 273.15$$

IDEAL GAS LAWS

P = pressure
V = volume (expressed in liters)
T = temperature (expressed in Kelvin)
Standard temperature and pressure (STP):
 $T = 0^\circ\text{C} = 273 \text{ K}$
 $P = 1 \text{ atm} = 760 \text{ mmHg} = 101.3 \text{ kPa}$

A. Boyle's law: $PV = \text{constant}$

$$P_1V_1 = P_2V_2$$

B. Charles' law: $\frac{V}{T} = \text{constant}$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

C. Gay-Lussac's law: $\frac{P}{T} = \text{constant}$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

D. Combined gas law: $\frac{PV}{T} = \text{constant}$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

E. Ideal gas law: $PV = nRT$
 $n =$ number of moles, $R =$ ideal gas constant
 $R = 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} = 8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}}$

F. Dalton's law of partial pressures:

$$\frac{P_{\text{gas}}}{P_{\text{total}}} = \frac{n_{\text{gas}}}{n_{\text{total}}}$$

$$\frac{P_{\text{gas}}}{P_{\text{total}}} = \frac{V_{\text{gas}}}{V_{\text{total}}}$$

$$\frac{P_{\text{gas}}}{P_{\text{total}}} = \frac{m_{\text{gas}}}{m_{\text{total}}}$$

$$P_{\text{total}} = P_{\text{gas}} + P_{\text{water}}$$

G. Graham's law of effusion: $\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$

H. Diffusion: $\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$

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$\Delta T_b = K_b m$, where ΔT_b is the amount the boiling point is lowered, K_b is a proportionality constant specific to the solvent, and m is the molality of the solution.

E. Boiling point elevation:

$$\Delta T_b = K_b m$$

ΔT_b is the amount the boiling point is elevated, K_b is a proportionality constant specific to the solvent, and m is the molality of the solution.

F. Osmotic pressure:

$$\pi = MRT$$

where π is the osmotic pressure, M is the molarity of the solution, R is the ideal gas constant, and T is the temperature in Kelvin.

CHEMICAL EQUILIBRIA

A. Equilibrium constant expression:

Method using molar concentration: For the reaction: $aA + bB \rightleftharpoons cC + dD$

$$\text{Equilibrium constant } K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Method using gas partial pressures: For the reaction: $aA(g) + bB(g) \rightleftharpoons cC(g) + dD(g)$

$$\text{Equilibrium constant } K_p = \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$$

B. Acid-base equilibrium:

1. Acid dissociation reactions: The reaction of a weak acid with H_2O .



Acid dissociation constant: Equilibrium constant for the reaction:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

The greater the K_a value, the stronger the acid.

2. Base dissociation reactions: The reaction of a weak base with H_2O .

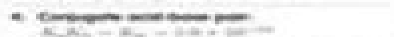


Base dissociation constant: Equilibrium constant for the reaction:

$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

The greater the K_b value, the stronger the base.

3. Dissociation of water: The self-ionization of water:



$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

4. Conjugate acid-base pair:

$$K_a K_b = K_w = 1.0 \times 10^{-14}$$

5. pH and pOH: pH measures $10^{-\text{pH}}$ molar H_3O^+ .

$$\text{Acids: pH} < 7 \text{ and pOH} > 7$$

$$\text{Bases: pH} > 7 \text{ and pOH} < 7$$

$$\text{Water: pH} = \text{pOH} = 7$$

$$\text{Calculating pH: } \text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$$

$$\text{Calculating pOH: } \text{pOH} = -\log_{10}[\text{OH}^-]$$

$$\text{For all aqueous solutions: } \text{pH} + \text{pOH} = 14$$

$$\text{For 25 } ^\circ\text{C: } \text{pH} = 7 \text{ and } \text{pOH} = 7$$

concentration with respect to H_2O , assuming the reaction rate is independent of $[\text{H}_2\text{O}]$.

THEORETICAL AND PERCENTAGE YIELD

Percentage yield = $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100$

THERMODYNAMICS

A. Specific heat:

$q = mc\Delta T$, where q is the heat absorbed by the object, m is the mass of the object, ΔT is the change in temperature, and c is the specific heat of the substance.

B. Enthalpy (ΔH): In kJ mol^{-1}

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

$\Delta H < 0$ for exothermic reactions
 $\Delta H > 0$ for endothermic reactions

C. Entropy (ΔS): In $\text{J mol}^{-1} \text{K}^{-1}$

$$\Delta S = S_{\text{products}} - S_{\text{reactants}}$$

D. Gibbs free energy (ΔG): In kJ mol^{-1}

$$\Delta G = \Delta H - T\Delta S$$

$\Delta G < 0$ when forward reaction is spontaneous
 $\Delta G > 0$ when reverse reaction is spontaneous

REDOX REACTIONS AND ELECTROCHEMISTRY

A. Standard reduction potential of a cell:

$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$$

B. Nernst equation:

$$E = E^\circ - \frac{RT}{nF} \ln Q$$

where E is the cell potential, Q is the reaction quotient, R is the gas constant, T is the temperature in Kelvin, n is the number of electrons transferred in the half-reaction, F (Faraday's constant) = 96485 C mol^{-1} , and $\ln$ is the natural logarithm.

C. Standard cell voltage at equilibrium:

$$E_{\text{cell}}^\circ = \frac{RT}{nF} \ln K$$

where n is the total number of electrons transferred in the redox reaction.

D. Thermodynamics:

$$\Delta G^\circ = -nFE_{\text{cell}}^\circ$$

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

$$\ln K_{\text{eq}} = \frac{nE_{\text{cell}}^\circ}{RT}$$

$$K_{\text{eq}} = e^{\frac{nE_{\text{cell}}^\circ}{RT}}$$

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Emma Harrison



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