Form 2

CHEMISTRY 1202

MAIN Final Exam – Spring 2011

Time allowed: **2 hours**

INSTRUCTIONS:

On "side one" of your scantron sheet:

1. Write your name and (the name of your CHEM 1202 lecturer in parentheses)

On "side two" of your scantron sheet:

1. In "NAME", write and bubble in your name, last name ***FIRST***;
2. In "IDENTIFICATION NUMBER", write and bubble in your LSU ID NUMBER;
3. In "GRADE OR EDUCATION", bubble in the FORM NUMBER **2**;
4. At the end of the exam, turn the scantron sheet to the exam proctors.

Useful constants and formulas:

$R = 8.314\ ^{J}/_{mol\,K}$ *or* $R = 0.0821\ ^{atm\,L}/_{mol\,K}$ $F = 96{,}485\ ^{C}/_{mol}$ *or* $^{J}/_{mol\,V}$ $K_w = 1.0 \times 10^{-14}$

Integrated Rate Laws:

$$[A]_t = -kt + [A]_0 \qquad \ln\frac{[A]_t}{[A]_0} = -kt \qquad \frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$

Arrhenius Eq^n: $\ln\frac{k_2}{k_1} = \frac{E_a}{R}\left(\frac{T_2 - T_1}{T_1 \bullet T_2}\right)$ *or* $\ln\frac{k_2}{k_1} = \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$

Nernst Eq^n: $E_{cell} = E^\circ_{cell} - \frac{0.0592}{n}\log Q$ *or* $E_{cell} = E^\circ_{cell} - \frac{RT}{nF}\ln Q$

Formula 19.16: $\Delta G = \Delta G^\circ + RT\ln Q$

Select standard reduction potentials:

$F_2(g) + 2e^- \rightarrow 2F^-(aq)$	$E^\circ_{red} = +2.87$ V
$AgBr(s) + e^- \rightarrow Ag(s) + Br^-(aq)$	$E^\circ_{red} = +0.095$ V
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	$E^\circ_{red} = +0.536$ V
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	$E^\circ_{red} = +0.337$ V
$2H^+(aq) + 2\,e^- \rightarrow H_2(g)$	**$E^\circ_{red} = 0.000$ V**
$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$	$E^\circ_{red} = -0.403$ V
$Co^{2+}(aq) + 2e^- \rightarrow Co(s)$	$E^\circ_{red} = -0.277$ V
$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$	$E^\circ_{red} = -1.66$ V
$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$	$E^\circ_{red} = -2.37$ V

1. The equilibrium constant K_p is 1.0×10^{-13} for the reaction

$$2\ CO_2(g) \leftrightarrows 2\ CO(g) + O_2(g)$$

For which of the following initial conditions will the reaction proceed to the **right**?

(i) $P_{CO2} = 1.0 \times 10^{-3}$ atm; $P_{CO}=1.0 \times 10^{-7}$ atm; $P_{O2} = 2.0 \times 10^{-3}$ atm

(ii) $P_{CO2} = 1.0 \times 10^{-1}$ atm; $P_{CO}=1.0 \times 10^{-5}$ atm; $P_{O2} = 2.0 \times 10^{-7}$ atm

(iii) $P_{CO2} = 1.0 \times 10^{-2}$ atm; $P_{CO}=1.0 \times 10^{-4}$ atm; $P_{O2} = 2.0 \times 10^{-6}$ atm

A. (i) only
B. (ii) only
C. (iii) only
D. (i) and (ii) only
E. (i) and (iii) only

2. What is the correct equilibrium constant expression for the reaction

$$CH_3COH(aq) + 2\ CH_3OH(aq) \leftrightarrows CH_3(OCH_3)_2H(aq) + H_2O(l)?$$

A. $K = \dfrac{[CH_3(OCH_3)_2H]}{[CH_3COH][CH_3OH]^2}$

B. $K = \dfrac{[CH_3(OCH_3)_2H]}{[CH_3COH][CH_3OH]}$

C. $K = \dfrac{[CH_3COH][CH_3OH]^2}{[CH_3(OCH_3)_2]}$

D. $K = \dfrac{[CH_3COH][CH_3OH]}{[CH_3(OCH_3)_2H]}$

E. $K = \dfrac{[CH_3COH][CH_3OH]}{[CH_3(OCH_3)_2H][H_2O]}$

3. The following reaction is exothermic:

$$2\ CH_3COH(g) + 5\ O_2(g) \leftrightarrows 4\ CO_2(g) + 4\ H_2O(g)$$

Which of the following "stresses" will shift the equilibrium to the **right**?

(i) addition of oxygen
(ii) decreasing temperature
(iii) removal of carbon dioxide

A. (i) only
B. (ii) only
C. (iii) only
D. (i) and (ii) only
E. all

4. Consider the following reaction:

$$C(s) + CO_2(g) \leftrightarrows 2\ CO(g); \qquad K_c = 1.9$$

What is the equilibrium concentration of CO(g) if initially $[CO_2] = 0.20$ M and $[CO] = 0.00$ M?

A. 0.039 M
B. 0.075 M
C. 0.25 M
D. 0.30 M
E. 0.60 M

5. Which of the following represent correct conjugate acid (left) / conjugate base (right) pair(s)?

	Conjugate acid	*Conjugate base*
(i)	H_2O	H_3O^+
(ii)	HCO_3	H_2CO_3
(iii)	HCO_3^-	CO_3^{2-}
(iv)	OH^-	O^{2-}

A. (ii) only
B. (i) & (ii) only
C. (iii) only
D. (iii) & (iv) only
E. (iv) only

6. The pH of a solution is 5.20. What is the $[H_3O^+]$ at this pH?

A. 5.5×10^{-3} M
B. 1.6×10^{-9} M
C. 2.0×10^{-5} M
D. 6.3×10^{-6} M
E. 4.8×10^{-4} M

7. Given the following acids and their associated pK_a values, which of the acids is the **strongest**?

A. HIO_3 ($pK_a = 0.77$)
B. HClO ($pK_a = 2.0$)
C. CH_3COOH ($pK_a = 4.7$)
D. NH_4^+ ($pK_a = 9.2$)
E. HF ($pK_a = 3.2$)

8. What is the **pH** of a 0.10 M hypothetical acid, HA, solution that has a $K_a = 2.0 \times 10^{-5}$?

$$HA\ (aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + A^-(aq)$$

A. pH = 1.85
B. pH = 2.35
C. pH = 2.85
D. pH = 4.35
E. pH = 5.70

9. Which of the following **salts** results in an aqueous solution with **basic** pH?

(i) NaCl
(ii) Na_2CO_3
(iii) $NaCH_3COO$
(iv) NH_4Br

A. (i) only
B. (ii) only
C. (iii) only
D. (iv) only
E. (ii) & (iii) only

10. A titration of 60.0 mL of 0.30 M pyridine, C_5H_5N, required 90.0 mL of 0.20 M HCl to reach the equivalence point. What is the pH at this equivalence point? $K_b(C_5H_5N) = 2.0 \times 10^{-9}$

A. pH = 2.86
B. pH = 3.11
C. pH = 4.80
D. pH = 9.20
E. pH = 11.00

11. Which of the following mixtures constitute a **buffer solution**?

(i) 0.1 M NaCl + 0.1 M HCl
(ii) 0.1 M CH_3COOH + 0.1 M $NaCH_3COO$
(iii) 0.1 M $NaHCO_3$ + 0.1 M NH_4Cl
(iv) 0.1 M HClO + 0.1 M NaClO

A. (i) only
B. (ii) only
C. (iii) only
D. (ii) & (iv) only
E. (i), (ii) & (iv) only

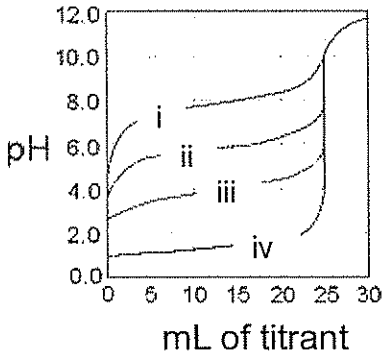
12. Consider a generic indicator, HIn, which is a very weak acid. The acid form of this indicator, HIn, is colorless and exits in equilibrium with its conjugate base form, In^-, which is red.

$$HIn(aq) + H_2O(l) \leftrightarrows In^-(aq) + H_3O^+(aq)$$

When excess HCl is added to the colorless solution containing the above indicator, _____.

A. the equilibrium shifts to the right and the solution turns red
B. the equilibrium shifts to the left and the solution remains colorless
C. the equilibrium shifts to the left and the solution turns red
D. the equilibrium shifts to the right and the solution remains colorless
E. the equilibrium doesn't shift in either direction and the solution remains colorless

13. Consider the figure with titrations curves for acids (i) – (iv).



Which of the following statements is **true**?

A. Acid (iv) is the weakest of the titrated acids.

B. Acid (ii) has a stronger conjugate base than acid (i) does.

C. Acid (iii) is stronger than acid (ii).

D. Acid (i) has a weaker conjugate base than acids (ii) and (iii) do.

E. Acid (ii) has a $pK_a \cong 4$.

14. What is the solubility of $Mg_3(AsO_4)_2(s)$ in water at 25 °C? $K_{sp}(Mg_3(AsO_4)_2) = 2.1 \times 10^{-20}$.

A. 2.7×10^{-5} M

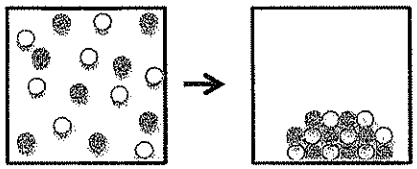
B. 4.5×10^{-5} M

C. 6.2×10^{-5} M

D. 7.7×10^{-5} M

E. 9.4×10^{-5} M

15. The spontaneous deposition process (gas → solid) shown below is an **exothermic** phase change.



Which of the following is **true**?

A. $\Delta S_{surr} < 0$ and $\Delta S_{sys} < 0$

B. $\Delta S_{surr} < 0$ and $\Delta S_{sys} > 0$

C. $\Delta S_{surr} > 0$ and $\Delta S_{sys} < 0$

D. $\Delta S_{surr} > 0$ and $\Delta S_{sys} = 0$

E. $\Delta S_{surr} = 0$ and $\Delta S_{sys} < 0$

16. For a particular process, the heat, q, is –22 kJ and the pressure-volume work, w, is 52 kJ. What conclusion(s) can be drawn for this process?

(i) $\Delta E = 30.$ kJ
(ii) $\Delta S_{surr} > 0$
(iii) the process is endothermic
(iv) the volume of the system decreases

A. (i) only
B. (ii) only
C. (i) and (ii) only
D. (i), (ii) and (iv) only
E. all statements are correct

17. Which species has the standard molar enthalpy of formation, $\Delta H_f°$, equal to zero?

A. $N_2(g)$
B. $NO(g)$
C. $NO_2(g)$
D. $HNO_2(aq)$
E. $NH_3(g)$

18. Use the ΔH values for reactions (1) and (2) to determine the value of ΔH_3 for equation (3).

(1) $2\ NO(g) \rightarrow N_2(g) + O_2(g)$ $\Delta H_1 = -180.$ kJ
(2) $2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$ $\Delta H_2 = -112$ kJ
(3) $N_2(g) + 2\ O_2(g) \rightarrow 2\ NO_2(g)$ $\Delta H_3 = ?$

A. $\Delta H_3 = +68$ kJ
B. $\Delta H_3 = +180.$ kJ
C. $\Delta H_3 = -112$ kJ
D. $\Delta H_3 = -292$ kJ
E. $\Delta H_3 = -404$ kJ

19. Consider the following reaction of combustion of ethane, C_2H_6:

$$C_2H_6(g) + {}^7/_2\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(l) \qquad \Delta H°_{rxn} = -1660 \text{ kJ}$$

What is the heat of the reaction measured at constant pressure when 0.178 mol of ethane is combusted?

A. –591 kJ
B. –295 kJ
C. –148 kJ
D. 17.8 kJ
E. 9,330 kJ

Use the following information to answer the next <u>two</u> questions

	2 PbS(s)	+	3 O_2(g)	⇆	2 PbO(s)	+	2 SO_2(g)
ΔH°_f **(kJ/mol)**	−100.		0		−217.3		−296.9
ΔG°_f **(kJ/mol)**	−98.7		0		−187.9		−300.4
S° **(J/mol K)**	91.2		205.0		68.7		248.5

20. **Using ΔG°_f values**, calculate ΔG°_{rxn} for the above reaction at 298 K?

A. +389.6 kJ
B. −389.6 kJ
C. −828.0 kJ
D. +779.2 kJ
(E.) −779.2 kJ

21. Assuming that ΔH°_f and S° values are independent of temperature, determine theoretically at what temperature, T, the above reaction becomes spontaneous at standard conditions.

A. T = 122 K
B. T = 278 K
(C.) T = 5082 K
D. Such a temperature doesn't exist; this process is always spontaneous.
E. Such a temperature doesn't exist; this process is always nonspontaneous.

22. What is the oxidation number of S in SO_4^{2-}?

A. 0 B. +2 C. +4 (D.) +6 E. +8

23. Consider the following redox reaction:

$$Zn(s) + 2\ H^+(aq) \leftrightarrows Zn^{2+}(aq) + H_2(g);\ E^\circ = -0.76\ V$$

For this reaction the following are **true** at standard conditions:

A. $\Delta G^\circ > 0$ and $K > 1$
B. $\Delta G^\circ < 0$ and $K > 1$
C. $\Delta G^\circ = 0$ and $K > 1$
(D.) $\Delta G^\circ > 0$ and $K < 1$
E. $\Delta G^\circ = 0$ and $K < 1$

24. Consider the following reaction:

$$H_2O(l) \leftrightarrows H^+(aq) + OH^-(aq) \qquad K_w = 2.916 \times 10^{-15} \text{ at } 10\ ^\circ C$$

What is the value of ΔG_{rxn} at 10 °C when $[H^+] = 5.4 \times 10^{-8} M$ and $[OH^-] = 5.4 \times 10^{-8} M$?

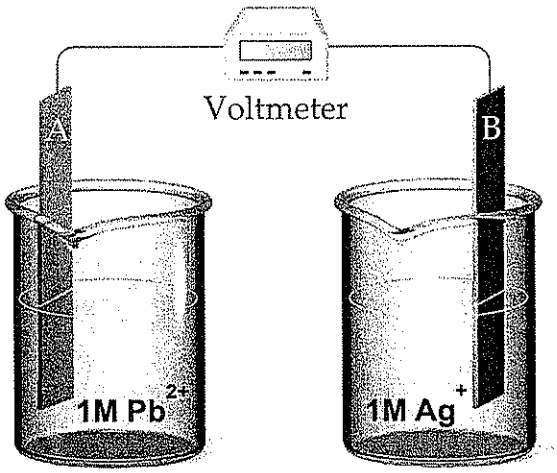
A. 0 kJ/mol
B. 2.8 kJ/mol
C. –39 kJ/mol
D. +39 kJ/mol
E. 79 kJ/mol

25. Assume that you want to construct a voltaic (galvanic) cell that uses the following two reactions:

$$Pb^{2+}(aq) + 2\ e^- \rightarrow Pb(s) \qquad E^\circ_{red} = -0.13\ V$$

$$Ag^+(aq) + e^- \rightarrow Ag(s) \qquad E^\circ_{red} = +0.80\ V$$

Beginning with the **incomplete** cell picture below, where electrode A is the anode and electrode B is the cathode, which of the following additions has to be made for the cell to generate a standard cell potential?



A. add salt bridge to enable the flow of ions from one half-cell to the other.
B. immerse Ag(s) electrode in the 1M $Pb^{2+}(aq)$ solution.
C. make sure electrode A is made of solid silver.
D. immerse Pb(s) electrode in the 1M $Ag^+(aq)$ solution.
E. remove the voltmeter

26. When the following redox reaction is balanced in **acidic** solution, the coefficients in front of O_2 and Fe^{3+} are x and y, respectively.

$$O_2(g) + Fe^{2+}(aq) \rightarrow H_2O(l) + Fe^{3+}(aq)$$

A. $x = 1$ and $y = 1$
(B) $x = 1$ and $y = 4$
C. $x = 2$ and $y = 4$
D. $x = 2$ and $y = 3$
E. $x = 4$ and $y = 4$

27. The following data were collected for the rate of disappearance of NO in the reaction:

$$2\ NO(g) + O_2\ (g) \rightarrow 2\ NO_2(g)$$

The initial rate experiment resulted in the following data:

Experiment Number	**[NO] (M)**	**[O_2] (M)**	**initial rate (M/s)**
1	0.0126	0.0125	1.41×10^{-2}
2	0.0252	0.0250	1.13×10^{-1}
3	0.0252	0.0125	5.64×10^{-2}

What is the rate law for the reaction?

A. Rate = $k[NO]^2$
B. Rate = $k[NO][O_2]$
C. Rate = $k[NO][O_2]^2$
D. Rate = $k[O_2]$
(E) Rate = $k[NO]^2[O_2]$

28. The rate constant for a reaction increases by a factor of 10 when the temperature is raised from 250K to 325K. What is the activation energy for this reaction?

(A) 20.7 kJ/mol
B. 24.8 kJ/mol
C. 28.5 kJ /mol
D. 33.2 kJ/mol
E. 38.1 kJ/mol

29. What is the rate law for the overall reaction

$$2\ NO_2(g) + F_2(g) \rightarrow 2\ NO_2F(g)$$

if the mechanism consists of these two elementary reactions:

$NO_2(g) + F_2(g) \rightarrow NO_2F(g) + F(g)$ Step 1, *slow*

$NO_2(g) + F(g) \rightarrow NO_2F(g)$ Step 2, *fast*

A. Rate = $k[NO_2]^2[F_2]$
B. Rate = $k[NO_2][F_2]$
C. Rate = $k[NO_2][F]$
D. Rate = $k[NO_2F][F]$
E. Rate = $k[NO_2F]$

30. The reaction $SO_2Cl_2(g) \rightarrow SO_2(g) + Cl_2(g)$ is a first order reaction. If the initial concentration of $SO_2Cl_2 = 5.00 \times 10^{-3}$ M and the concentration of SO_2Cl_2 after 1.000×10^3 seconds is 4.89×10^{-3} M, what is the rate constant for this reaction?

A. $2.93 \times 10^{-6}\ s^{-1}$
B. $4.08 \times 10^{-6}\ s^{-1}$
C. $8.19 \times 10^{-5}\ s^{-1}$
D. $2.22 \times 10^{-5}\ s^{-1}$
E. $5.66 \times 10^{-5}\ s^{-1}$

31. At 298 K, what is the cell potential for the reaction:

$$2\ Al^{3+}(aq) + 6\ I^-(aq) \leftrightarrows 2\ Al(s) + 3\ I_2(s)$$

If the initial concentrations are $[Al^{3+}] = 1.2 \times 10^{-3}$ M and $[I^-] = 2.8 \times 10^{-4}$ M?

A. -2.53 V
B. -2.46 V
C. -2.20 V
D. -1.86 V
E. -1.61 V

32. Consider the following reaction:

$$C_2H_4(g) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 2\,H_2O(g)$$

Which of the following equations is correct for this reaction?

A. Rate $= -\frac{2}{1}\frac{\Delta[C_2H_4]}{\Delta t} = -\frac{2}{3}\frac{\Delta[O_2]}{\Delta t} = \frac{\Delta[CO_2]}{\Delta t} = \frac{\Delta[H_2O]}{\Delta t}$

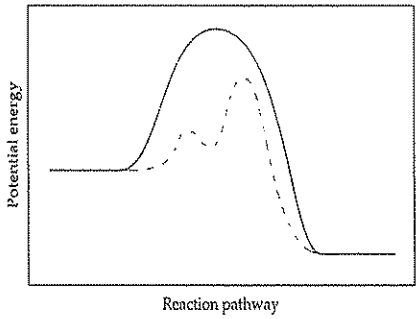
B. Rate $= \frac{\Delta[C_2H_4]}{\Delta t} = -\frac{1}{3}\frac{\Delta[O_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[CO_2]}{\Delta t} = \frac{\Delta[H_2O]}{\Delta t}$

C. Rate $= \frac{\Delta[C_2H_4]}{\Delta t} = 3\frac{\Delta[O_2]}{\Delta t} = 2\frac{\Delta[CO_2]}{\Delta t} = 2\frac{\Delta[H_2O]}{\Delta t}$

D. Rate $= -\frac{\Delta[C_2H_4]}{\Delta t} = -\frac{\Delta[O_2]^3}{\Delta t} = \frac{\Delta[CO_2]^2}{\Delta t} = \frac{\Delta[H_2O]^2}{\Delta t}$

(E) Rate $= -\frac{\Delta[C_2H_4]}{\Delta t} = -\frac{1}{3}\frac{\Delta[O_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[CO_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[H_2O]}{\Delta t}$

33. The following energy profile diagram is for the reaction of decomposition of hydrogen peroxide. Normally, this reaction is rather slow; however, an addition of a catalyst, such as HBr(aq), speeds it up.



Which of the statements below is **true**?

A. The solid line corresponds to the catalyzed reaction with a lower activation energy.
(B) The solid line corresponds to the uncatalyzed reaction with a higher activation energy.
C. The dashed line corresponds to the uncatalyzed reaction with a lower activation energy.
D. The dashed line corresponds to the catalyzed reaction with a higher activation energy.
E. The solid line corresponds to an uncatalyzed reaction with a lower activation energy.

Make sure to encode you ID number, name and form number 2 on the scantron sheet

WebElements: the periodic table on the world-wide web

http://www.webelements.com/

1 IA	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
hydrogen 1 H 1.0079																	helium 2 He 4.0026
lithium 3 Li 6.941	beryllium 4 Be 9.0122											boron 5 B 10.811	carbon 6 C 12.011	nitrogen 7 N 14.007	oxygen 8 O 15.999	fluorine 9 F 18.998	neon 10 Ne 20.180
sodium 11 Na 22.990	magnesium 12 Mg 24.305											aluminium 13 Al 26.982	silicon 14 Si 28.086	phosphorus 15 P 30.974	sulfur 16 S 32.065	chlorine 17 Cl 35.453	argon 18 Ar 39.948
potassium 19 K 39.098	calcium 20 Ca 40.078	scandium 21 Sc 44.956	titanium 22 Ti 47.867	vanadium 23 V 50.942	chromium 24 Cr 51.996	manganese 25 Mn 54.938	iron 26 Fe 55.845	cobalt 27 Co 58.933	nickel 28 Ni 58.693	copper 29 Cu 63.546	zinc 30 Zn 65.39	gallium 31 Ga 69.723	germanium 32 Ge 72.61	arsenic 33 As 74.922	selenium 34 Se 78.96	bromine 35 Br 79.904	krypton 36 Kr 83.80
rubidium 37 Rb 85.468	strontium 38 Sr 87.62	yttrium 39 Y 88.906	zirconium 40 Zr 91.224	niobium 41 Nb 92.906	molybdenum 42 Mo 95.94	technetium 43 Tc [98]	ruthenium 44 Ru 101.07	rhodium 45 Rh 102.91	palladium 46 Pd 106.42	silver 47 Ag 107.87	cadmium 48 Cd 112.41	indium 49 In 114.82	tin 50 Sn 118.71	antimony 51 Sb 121.76	tellurium 52 Te 127.60	iodine 53 I 126.90	xenon 54 Xe 131.29
caesium 55 Cs 132.91	barium 56 Ba 137.33	57-70 *	lutetium 71 Lu 174.97	hafnium 72 Hf 178.49	tantalum 73 Ta 180.95	tungsten 74 W 183.84	rhenium 75 Re 186.21	osmium 76 Os 190.23	iridium 77 Ir 192.22	platinum 78 Pt 195.08	gold 79 Au 196.97	mercury 80 Hg 200.59	thallium 81 Tl 204.38	lead 82 Pb 207.2	bismuth 83 Bi 208.98	polonium 84 Po [209]	astatine 85 At [210]	radon 86 Rn [222]
francium 87 Fr [223]	radium 88 Ra [226]	89-102 **	lawrencium 103 Lr [262]	rutherfordium 104 Rf [261]	dubnium 105 Db [262]	seaborgium 106 Sg [266]	bohrium 107 Bh [264]	hassium 108 Hs [269]	meitnerium 109 Mt [268]	darmstadtium 110 Ds [281]	roentgenium 111 Rg [272]	ununbium 112 Uub [285]	ununtrium 113 Uut [284]	ununquadium 114 Uuq [289]	ununpentium 115 Uup [288]	ununhexium 116 Uuh [292]	

*lanthanoids	lanthanum 57 La 138.91	cerium 58 Ce 140.12	praseodymium 59 Pr 140.91	neodymium 60 Nd 144.24	promethium 61 Pm [145]	samarium 62 Sm 150.36	europium 63 Eu 151.96	gadolinium 64 Gd 157.25	terbium 65 Tb 158.93	dysprosium 66 Dy 162.50	holmium 67 Ho 164.93	erbium 68 Er 167.26	thulium 69 Tm 168.93	ytterbium 70 Yb 173.04
**actinoids	actinium 89 Ac [227]	thorium 90 Th 232.04	protactinium 91 Pa 231.04	uranium 92 U 238.03	neptunium 93 Np [237]	plutonium 94 Pu [244]	americium 95 Am [243]	curium 96 Cm [247]	berkelium 97 Bk [247]	californium 98 Cf [251]	einsteinium 99 Es [252]	fermium 100 Fm [257]	mendelevium 101 Md [258]	nobelium 102 No [259]

Symbols and names: the symbols and names of the elements, and their spellings are those recommended by the International Union of Pure and Applied Chemistry (IUPAC - http://www.iupac.org/). Names have yet to be proposed for the most recently discovered elements 111–112 and 114 so those used here are IUPAC's temporary systematic names. In the USA and some other countries, the spellings aluminum and cesium are normal while in the UK and elsewhere the common spelling is sulphur.

Group labels: the numeric system (1–18) used here is the current IUPAC convention.

Atomic weights (mean relative masses): Apart from the heaviest elements, these are the IUPAC 2001 values and given to 5 significant figures. Elements for which the atomic weight is given within square brackets have no stable nuclides and are represented by the element's longest lived isotope.